Chapter 4

Reactions in Aqueous Solution

Solutions

- Solutions are defined as homogeneous mixtures of two or more pure substances.
- The *solvent* is present in greatest abundance.
- All other substances are *solutes*.

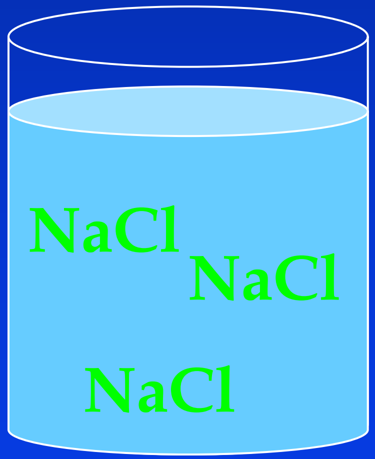
Parts of Solutions

- Solution- homogeneous mixture.
- Solute- what gets dissolved.
- Solvent- what does the dissolving.
- Soluble- Can be dissolved.
- Miscible- liquids dissolve in each other.

Solution Components

Solute – the substance that is dissolved; usually present in lesser quantities than the solvent

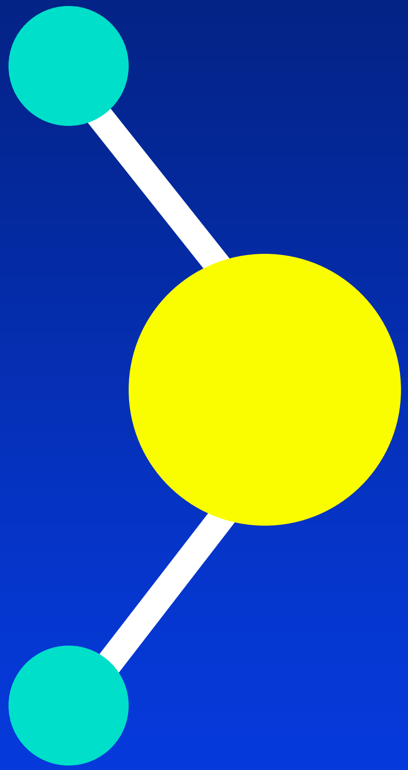
Solvent - the substance that does the dissolving (examples: water, hexane, methanol)



Making Solutions

Aqueous solutions

- Dissolved in water.
- Water is a good solvent because the molecules are polar.
- The oxygen atoms have a partial negative charge.
- The hydrogen atoms have a partial positive charge.
- The angle is 105^{o}C.

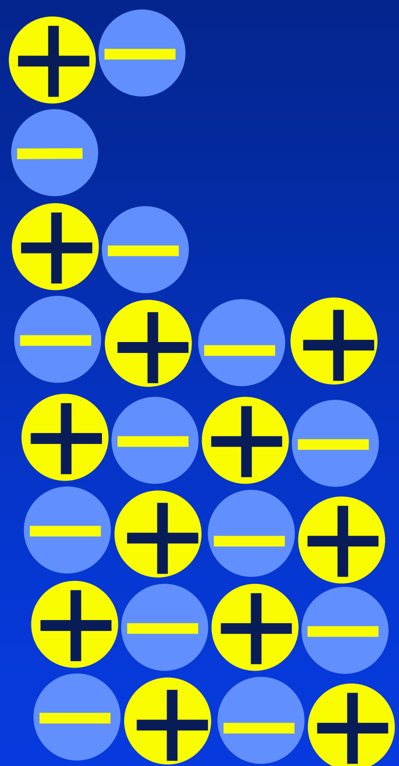


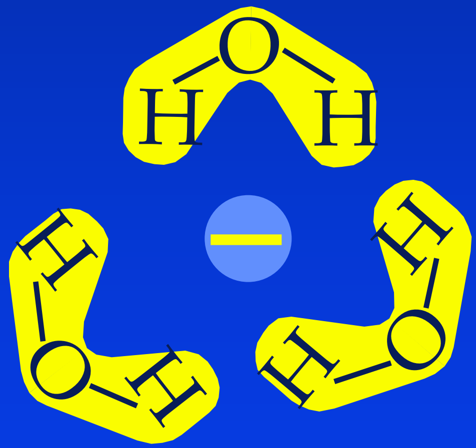
Hydration

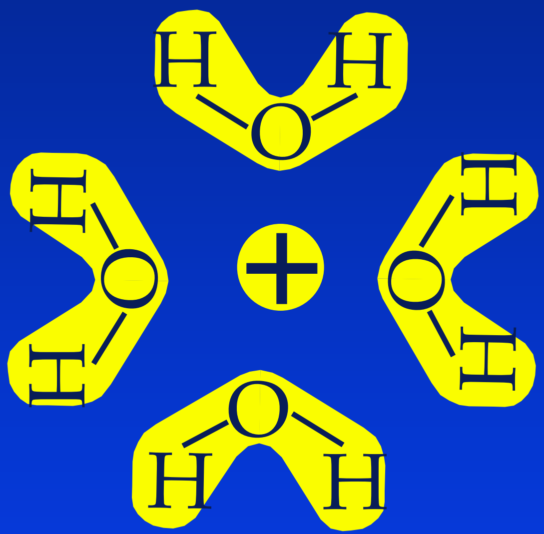
- The process of breaking the ions of salts apart.
- Ions have charges and are attracted to the opposite charges on the water molecules.

How Ionic solids dissolve

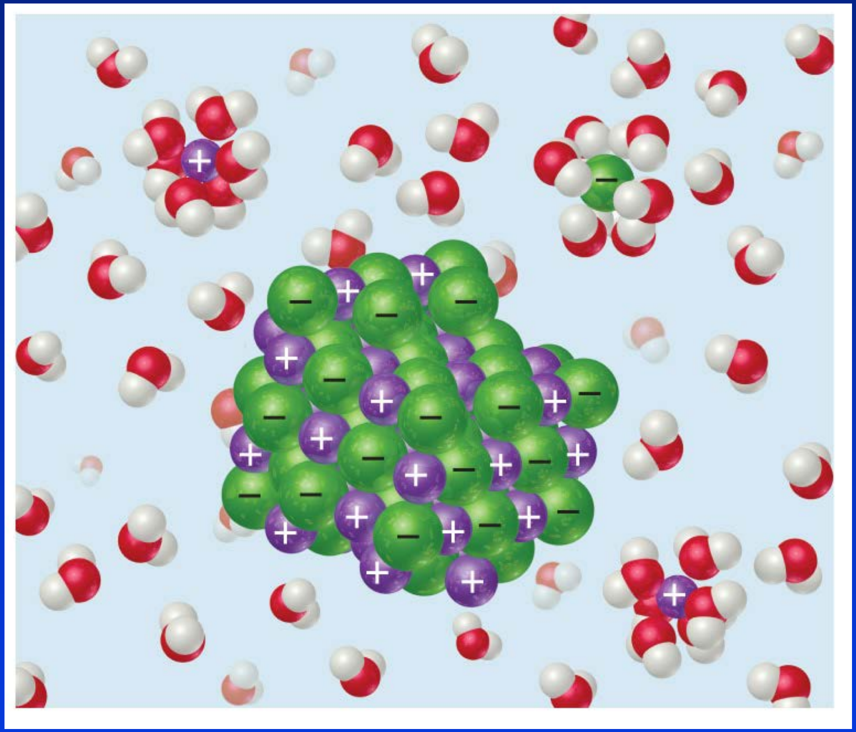
Click here for Animation





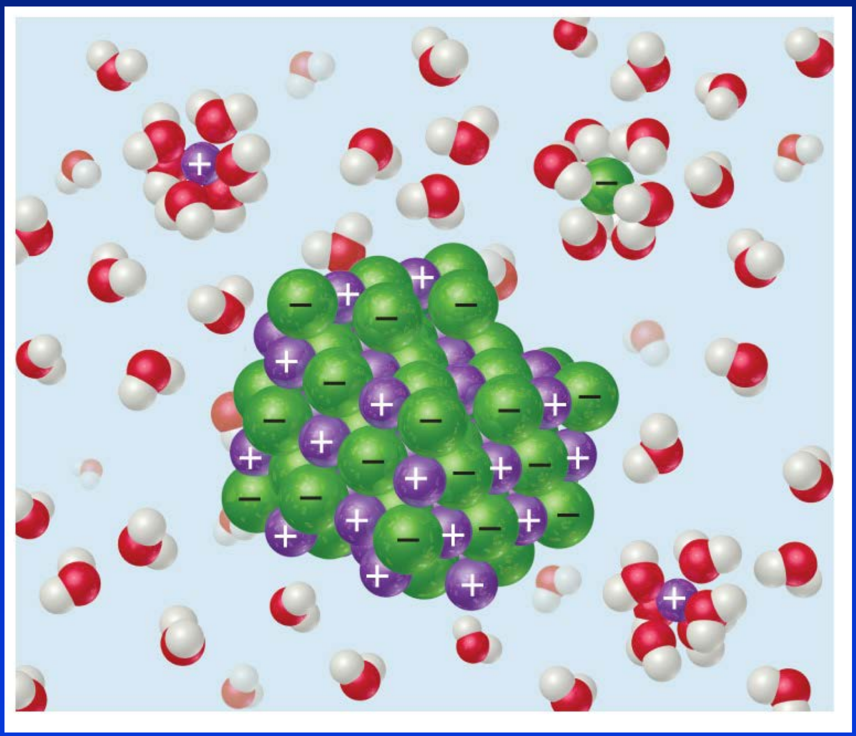


Dissociation



- When an ionic substance dissolves in water, the solvent pulls the individual ions from the crystal and solvates them.
- This process is called *dissociation*.

Dissociation



- An *electrolyte* is a substances that dissociates into ions when dissolved in water.

Solubility

- How much of a substance will dissolve in a given amount of water.
- Usually g/100 mL
- Varies greatly, but if they do dissolve the ions are separated,
- and they can move around.
- Water can also dissolve non-ionic compounds if they have polar bonds.

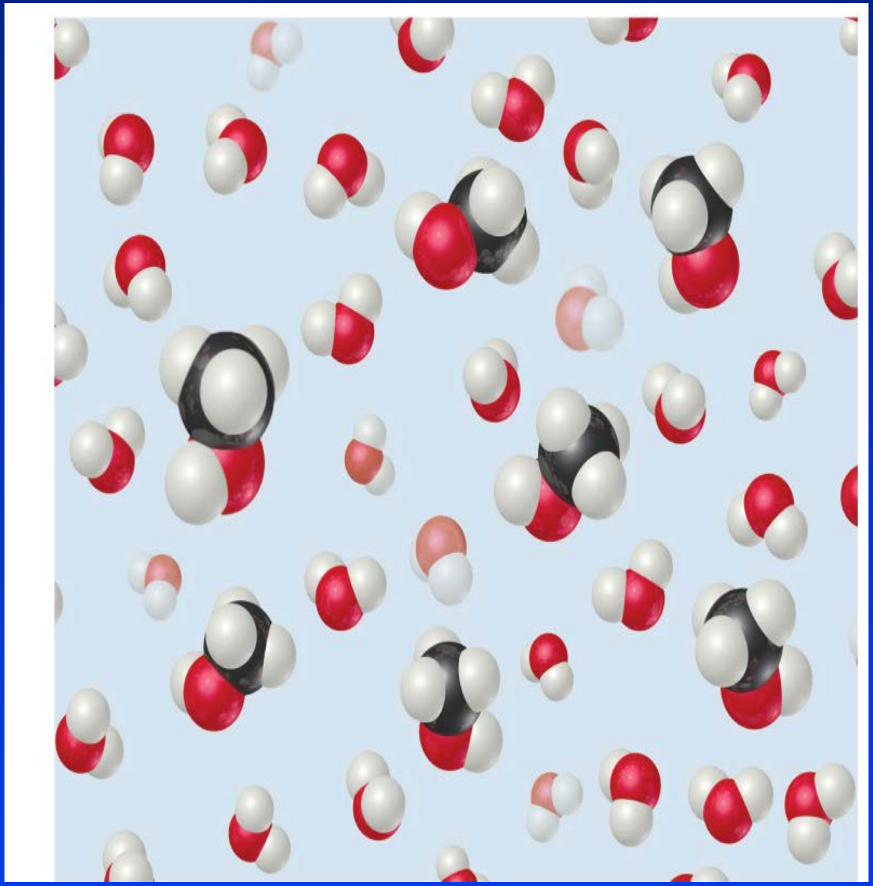
Electrolytes

- Electricity is moving charges.
- The ions that are dissolved can move.
- Solutions of ionic compounds can conduct electricity.
- Electrolytes.
- Solutions are classified three ways.

Types of solutions

- Strong electrolytes- completely dissociate (fall apart into ions).
  – Many ions- Conduct well.
- Weak electrolytes- Partially fall apart into ions.
  – Few ions -Conduct electricity slightly.
- Non-electrolytes- Don’t fall apart.
  – No ions- Don’t conduct.

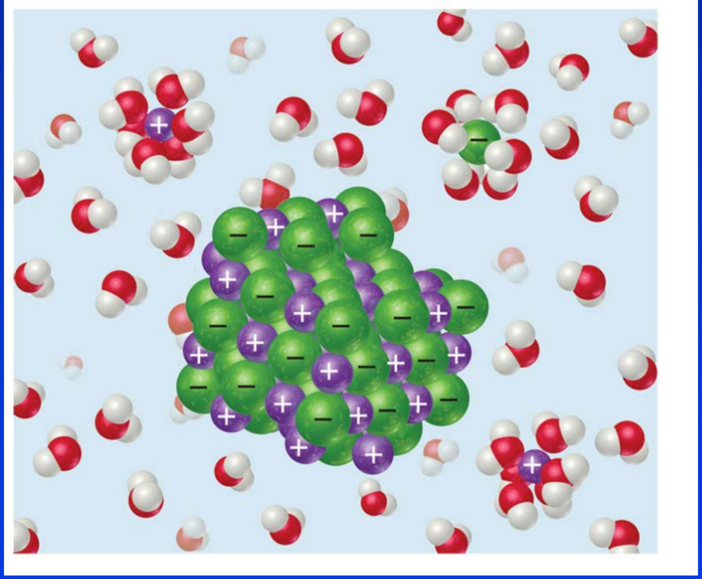
Electrolytes



- An electrolyte is a substances that dissociates into ions when dissolved in water.
- A nonelectrolyte may dissolve in water, but it does not dissociate into ions when it does so.

Electrolytes and Nonelectrolytes

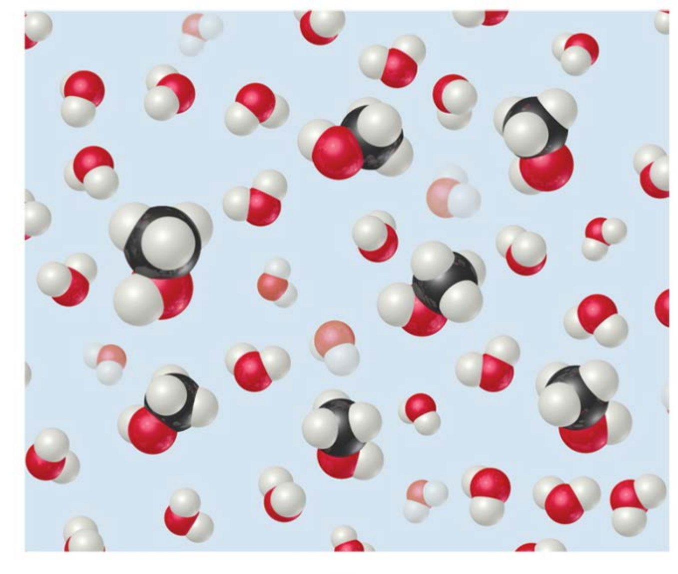
	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids Weak bases	All other compounds



Soluble ionic compounds tend to be electrolytes.

Electrolytes and Nonelectrolytes

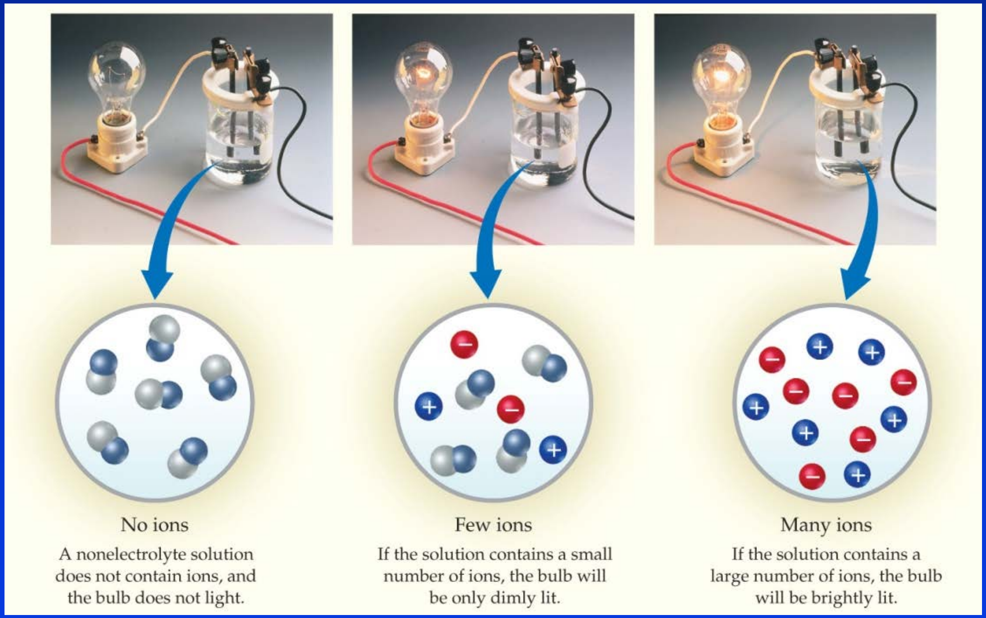
	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids Weak bases	All other compounds



Molecular compounds tend to be nonelectrolytes, except for acids and bases.

Electrolytes

- A strong electrolyte dissociates completely when dissolved in water.
- A weak electrolyte only dissociates partially when dissolved in water.



Strong Electrolytes Are…

- Strong acids
- Strong bases

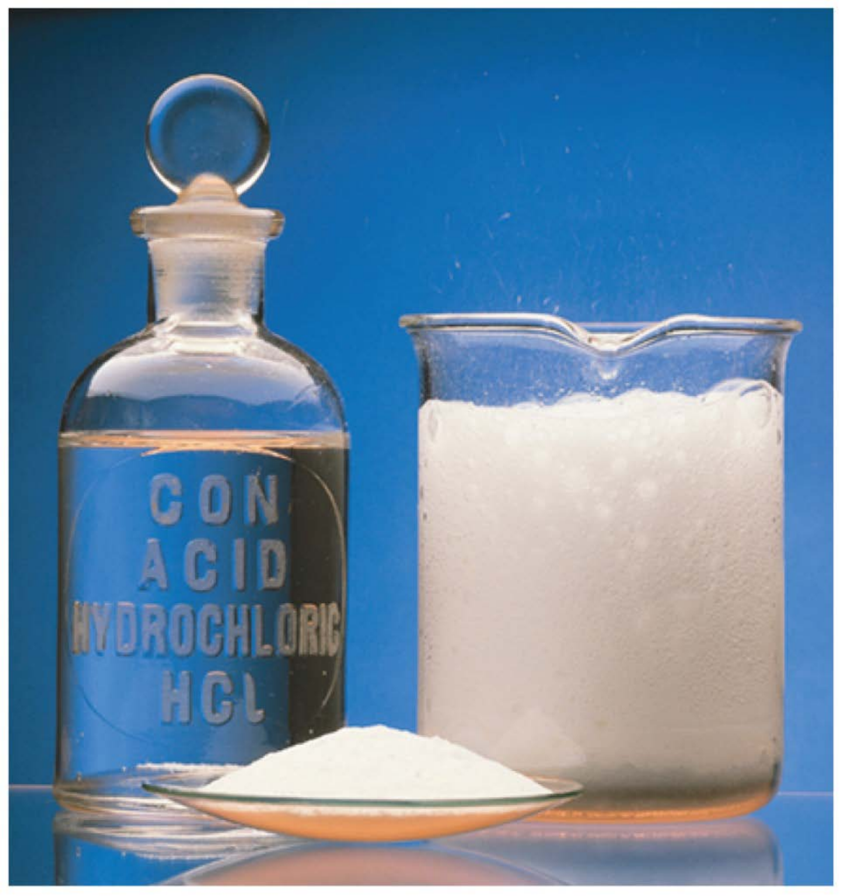
Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)
Hydrobromic, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic, HI	
Chloric, $HClO_3$	
Perchloric, $HClO_4$	
Nitric, HNO_3	
Sulfuric, H_2SO_4	

Strong Electrolytes Are…

- Strong acids
- Strong bases
- Soluble ionic salts

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of the alkali metal cations, and NH_4^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}

Acids



- Arrhenius defined acids as substances that increase the concentration of H^+ when dissolved in water.
- Brønsted and Lowry defined them as proton donors.

Acids

There are only seven strong acids:

- Hydrochloric (HCl)
- Hydrobromic (HBr)
- Hydroiodic (HI)
- Nitric (HNO_3)
- Sulfuric (H_2SO_4)
- Chloric ($HClO_3$)
- Perchloric ($HClO_4$)

Strong and Weak Acids

Strong acids are acids that are completely ionized in water and are good conductors of electricity

Introduction to Acids Clip

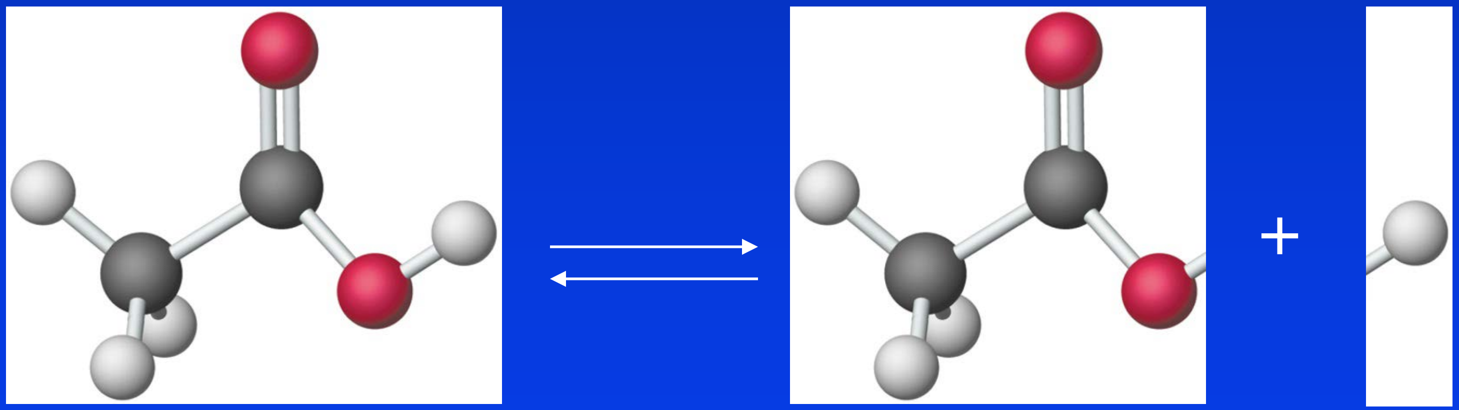
Weak acids are acids in which only some of the molecules are ionized in water; the rest remain as intact molecules

The dissociation of a weak acid in solution is written using a double arrow to indicate that the dissociation does not go to completion

Limited Ionization

Organic acids such as acetic acid are typical weak acids that have limited ionization in solution

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^{-}(aq) + H^{+}(aq)$$



Bases

- Arrhenius defined bases as substances that increase the concentration of OH^- when dissolved in water.
- Brønsted and Lowry defined them as proton acceptors.

Bases

The strong bases are the soluble metal salts of hydroxide ion:

- Alkali metals
- Calcium
- Strontium
- Barium

Strong and Weak Bases

- **Strong bases** are ionic hydroxides that completely ionize in water - good conductors of electricity
- **Weak bases** are substances that act as bases but remain mostly molecular at equilibrium in water
- The dissociation of a weak base in solution is written using a double arrow to indicate that the dissociation does not go to completion

Ammonia, NH_3, is a weak base

Types of solutions

- Acids- form H^+ ions when dissolved.
- Strong acids fall apart completely.
  - many ions
- Memorize this list
  H_2SO_4 HNO_3 HCl HBr HI $HClO_4$
- Weak acids- don’ dissociate completely.
- Bases - form OH^- ions when dissolved.
- Strong bases- many ions.
  - KOH NaOH

Concentrations

- Substances enter into chemical reactions according to certain molar **ratios**.
- Volumes of solutions are more convenient to measure than their masses.

Molar concentration expresses the amount of solute in one liter of solution

$$molarity\,(M) = \frac{mol\ solute}{L\ of\ solution}$$

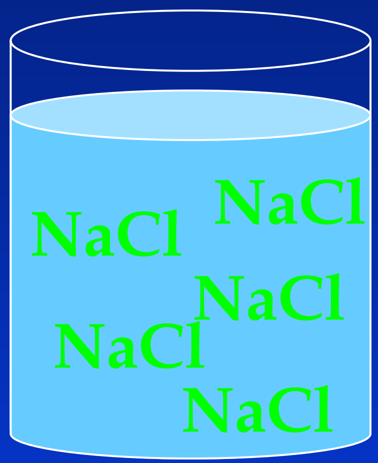
Examples of Concentration

Dilute solution

e.g., 0.1 M NaCl solution

(0.05 mol NaCl dissolved in 500 mL of water)

Concentrated solution



e.g., a 5 M NaCl solution

(5 mol NaCl dissolved in 1.0 L of water)

EOS

Molarity

- Two solutions can contain the same compounds but be quite different because the proportions of those compounds are different.
- Molarity is one way to measure the concentration of a solution.

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Measuring Solutions

- Concentration- how much is dissolved.
- Molarity = $\dfrac{\text{Moles of solute}}{\text{Liters of solution}}$
- abbreviated M
- 1 M = 1 mol solute / 1 liter solution
- Calculate the molarity of a solution with 34.6 g of NaCl dissolved in 125 mL of solution.

Molarity

- How many grams of HCl would be required to make 50.0 mL of a 2.7 M solution?
- What would the concentration be if you used 27g of $CaCl_2$ to make 500. mL of solution?
- What is the concentration of each ion?

Molarity

- Calculate the concentration of a solution made by dissolving 45.6 g of $Fe_2(SO_4)_3$ to 475 mL.
- What is the concentration of each ion?

Making solutions

- Describe how to make 100.0 mL of a 1.0 M $K_2Cr_2O_4$ solution.
- Describe how to make 250. mL of an 2.0 M copper (II) sulfate dihydrate solution.

Dilution

- One can also dilute a more concentrated solution by
  - Using a pipet to deliver a volume of the solution to a new volumetric flask, and
  - Adding solvent to the line on the neck of the new flask.

Dilution

The molarity of the new solution can be determined from the equation

$$M_c \times V_c = M_d \times V_d,$$

where M_c and M_d are the molarity of the concentrated and dilute solutions, respectively, and V_c and V_d are the volumes of the two solutions.

Dilution

- Adding more solvent to a known solution.
- The moles of solute stay the same.
- moles = M x L
- $M_1 V_1 = M_2 V_2$
- moles = moles
- Stock solution is a solution of known concentration used to make more dilute solutions

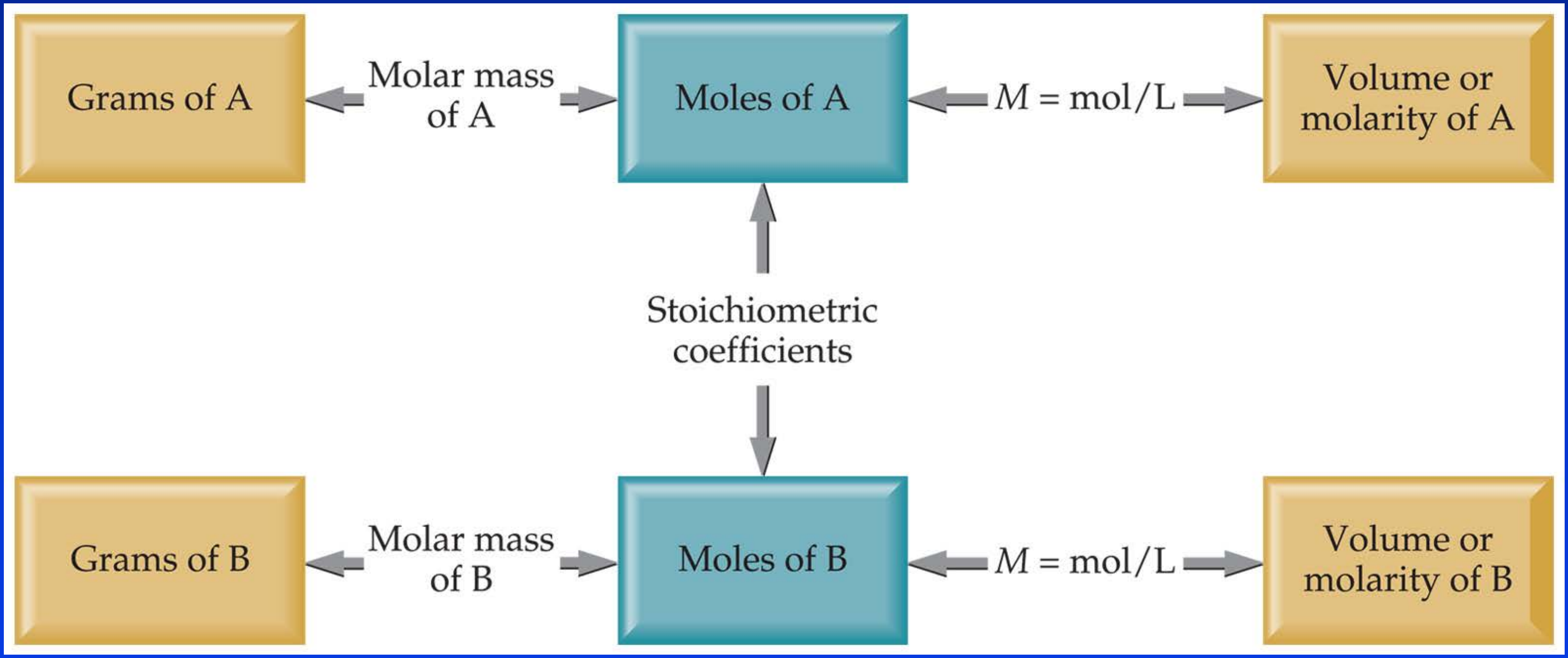
Dilution

- What volume of a 1.7 M solutions is needed to make 250 mL of a 0.50 M solution?
- 18.5 mL of 2.3 M HCl is added to 250 mL of water. What is the concentration of the solution?
- 18.5 mL of 2.3 M HCl is diluted to 250 mL with water. What is the concentration of the solution?

Dilution

- You have a 4.0 M stock solution. Describe how to make 1.0L of a 0.75 M solution.
- 25 mL 0.67 M of H_2SO_4 is added to 35 mL of 0.40 M $CaCl_2$. What mass $CaSO_4$ is formed?

Using Molarities in Stoichiometric Calculations



Chemical Reactions

All chemical reactions

- have two parts
- Reactants - the substances you start with
- Products- the substances you end up with
- The reactants turn into the products.
- Reactants $\rightarrow$ Products

Symbols used in equations

- the arrow separates the reactants from the products
- Read “reacts to form”
- Yields
- The plus sign = “and”
- (s) after the formula -solid
- (g) after the formula -gas
- (l) after the formula -liquid

Symbols used in equations

- (aq) after the formula - dissolved in water, an aqueous solution.
- ↑ used after a product indicates a gas (same as (g)) (products only)
- ↓ used after a product indicates a solid (same as (s)) (products only)

Symbols used in equations

- $\rightleftharpoons$ indicates a reversible reaction (More later)
- $\xrightarrow{\Delta}$, $\xrightarrow{\textbf{heat}}$ shows that heat is supplied to the reaction
- $\xrightarrow{\textbf{Pt}}$ is used to indicate a catalyst used in this case, platinum.

What is a catalyst?

- A substance that speeds up a reaction without being changed by the reaction.
- A catalyst lowers the activation energy
- Activation Energy- minimum amount of energy required by reacting particles in order to form the activated complex and lead to a reaction

Types of Reactions

1 Precipitation reactions

- When aqueous solutions of ionic compounds are poured together a solid forms.
- A solid that forms from mixed solutions is a precipitate
- If you’ re not a part of the solution, your part of the precipitate

Precipitation Reactions

When one mixes ions that form compounds that are insoluble (as could be predicted by the solubility guidelines), a precipitate is formed.

Precipitation reactions

$$NaOH_{(aq)} + FeCl_{3(aq)} \rightarrow NaCl_{(aq)} + Fe(OH)_{3(s)}$$

is really…

$$Na^{+}_{(aq)} + OH^{-}_{(aq)} + Fe^{+3} + Cl^{-}_{(aq)} \rightarrow$$
$$Na^{+}_{(aq)} + Cl^{-}_{(aq)} + Fe(OH)_{3(s)}$$

So all that really happens is…

$$OH^{-}_{(aq)} + Fe^{+3} \rightarrow Fe(OH)_{3(s)}$$

Precipitation reaction

- We can predict the products
- Can only be certain by experimenting
- The anion and cation switch partners
- $AgNO_3(aq) + KCl(aq) \rightarrow$
- $Zn(NO_3)_2(aq) + BaCr_2O_7(aq) \rightarrow$
- $CdCl_2(aq) + Na_2S(aq) \rightarrow$

Precipitations Reactions

- Will only happen if one of the products
  - doesn’t dissolve in water and forms a solid
  - or is a gas that bubbles out.
  - or is a covalent compound usually water.
- Polyatomic ions don’t change from side to side

Precipitations Reactions

- Only happen if one of the products is insoluble
- Otherwise all the ions stay in solution- nothing has happened.
- Need to memorize the rules for solubility (pg 145)

Applications of Precipitation Reactions

Reaction in Aqueous Solution	Application
$Al^{3+}(aq) + 3\,OH^-(aq) \longrightarrow Al(OH)_3(s)$	Water purification. (The gelatinous precipitate carries down suspended matter.)
$Al^{3+}(aq) + PO_4{}^{3-}(aq) \longrightarrow AlPO_4(s)$	Removal of phosphates from wastewater in sewage treatment.
$Mg^{2+}(aq) + 2\,OH^-(aq) \longrightarrow Mg(OH)_2(s)$	Precipitation of magnesium ion from seawater. (First step in the Dow process for extracting magnesium from seawater.)
$Ag^+(aq) + Br^-(aq) \longrightarrow AgBr(s)$	Preparation of AgBr for use in photographic film.
$Zn^{2+}(aq) + SO_4{}^{2-}(aq) + Ba^{2+}(aq) + S^{2-}(aq) \longrightarrow ZnS(s) + BaSO_4(s)$	Production of *lithopone*, a mixture used as a white pigment in both water paints and oil paints.
$H_3PO_4(aq) + Ca(OH)_2(aq) \longrightarrow CaHPO_4 \cdot 2\,H_2O(s)$	Preparation of calcium hydrogen phosphate dihydrate, used as a polishing agent in toothpastes.

Solubility Rules

1 All nitrates are soluble

2 Alkali metals ions and NH_4^+ ions are soluble

3 Halides are soluble except Ag^+, Pb^{+2}, and Hg_2^{+2}

4 Most sulfates are soluble, except Pb^{+2}, Ba^{+2}, Hg^{+2}, and Ca^{+2}

Solubility Rules

5 Most hydroxides are slightly soluble (insoluble) except NaOH and KOH

6 Sulfides, carbonates, chromates, and phosphates are insoluble

* Lower number rules supersede so Na_2S is soluble

Ionic Compounds and acids

- Fall apart into ions when they dissolve
- That's why they conduct electricity when dissolved.
- So when we write them as (aq) they are really separated
- $NaCl(aq)$ is really $Na^+(aq)$ and $Cl^-(aq)$
- K_2SO_4 (aq) is really $K^+(aq)$ and $SO_4^{2-}(aq)$

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Reactions in aqueous solutions

- Many reactions happen in solution
- Makes it so the ions separate so they can interact.
- Solids, liquids, and gases are not separated, only aqueous

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Three Types of Equations

- Molecular Equation- written as whole formulas, not the ions.

$$K_2CrO_{4(aq)} + Ba(NO_3)_{2(aq)} \rightarrow$$

- Complete Ionic equation show dissolved electrolytes as the ions.

$$2K^+ + CrO_4^{-2} + Ba^{+2} + 2\ NO_3^- \rightarrow$$
$$BaCrO_{4(s)} + 2K^+ + 2\ NO_3^-$$

- Spectator ions are those that don't react.

Three Type of Equations

- Net Ionic equations show only those ions that react, not the spectator ions

$$Ba^{+2} + CrO_4^{-2} \rightarrow BaCrO_{4(s)}$$

- Write the three types of equations for the reactions when these solutions are mixed.
  - iron (III) sulfate and potassium sulfide Lead (II) nitrate and sulfuric acid.

Writing Net Ionic Equations

1. Write a balanced molecular equation.
2. Dissociate all strong electrolytes.
3. Cross out anything that remains unchanged from the left side to the right side of the equation.
4. Write the net ionic equation with the species that remain.

Molecular Equation

The molecular equation lists the reactants and products in their molecular form.

$$AgNO_3\,(aq) + KCl\,(aq) \longrightarrow AgCl\,(s) + KNO_3\,(aq)$$

Ionic Equation

- In the ionic equation all strong electrolytes (strong acids, strong bases, and soluble ionic salts) are dissociated into their ions.
- This more accurately reflects the species that are found in the reaction mixture.

$$Ag^{+}\,(aq) + NO_3^{-}\,(aq) + K^{+}\,(aq) + Cl^{-}\,(aq) \longrightarrow$$
$$AgCl\,(s) + K^{+}\,(aq) + NO_3^{-}\,(aq)$$

Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.
- The only things left in the equation are those things that change (i.e., react) during the course of the reaction.
- Those things that didn't change (and were deleted from the net ionic equation) are called spectator ions.

$$Ag^{+}(aq) + \cancel{NO_3^{-}(aq)} + \cancel{K^{+}(aq)} + Cl^{-}(aq) \longrightarrow$$
$$AgCl\,(s) + \cancel{K^{+}(aq)} + \cancel{NO_3^{-}(aq)}$$

Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.

$$Ag^+(aq) + \cancel{NO_3^-}(aq) + \cancel{K^+}(aq) + Cl^-(aq) \longrightarrow$$
$$AgCl\,(s) + \cancel{K^+}(aq) + \cancel{NO_3^-}(aq)$$

Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.
- The only things left in the equation are those things that change (i.e., react) during the course of the reaction.

$$Ag^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl\,(s)$$

Gas-Forming Reactions

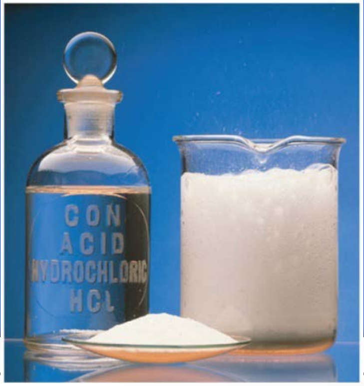
Some acid/base reactions form gases as products

Anion	Reaction with H^+
HCO_3^-	$HCO_3^- + H^+ \longrightarrow CO_2(g) + H_2O(l)$
CO_3^{2-}	$CO_3^{2-} + 2\,H^+ \longrightarrow CO_2(g) + H_2O(l)$
HSO_3^-	$HSO_3^- + H^+ \longrightarrow SO_2(g) + H_2O(l)$
SO_3^{2-}	$SO_3^{2-} + 2\,H^+ \longrightarrow SO_2(g) + H_2O(l)$
HS^-	$HS^- + H^+ \longrightarrow H_2S(g)$
S^{2-}	$S^{2-} + 2\,H^+ \longrightarrow H_2S(g)$

Gas-Forming Reactions

- Some metathesis reactions do not give the product expected.
- In this reaction, the expected product (H_2CO_3) decomposes to give a gaseous product (CO_2).

$$CaCO_3\,(s) + HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$



Gas-Forming Reactions

When a carbonate or bicarbonate reacts with an acid, the products are a salt, carbon dioxide, and water.

$$CaCO_3\ (s) + HCl\ (aq) \longrightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

$$NaHCO_3\ (aq) + HBr\ (aq) \longrightarrow NaBr\ (aq) + CO_2\ (g) + H_2O\ (l)$$

Gas-Forming Reactions

Similarly, when a sulfite reacts with an acid, the products are a salt, sulfur dioxide, and water.

$$SrSO_3\,(s) + 2\,HI\,(aq) \longrightarrow SrI_2\,(aq) + SO_2\,(g) + H_2O\,(l)$$

Gas-Forming Reactions

- This reaction gives the predicted product, but you had better carry it out in the hood, or you will be *very* unpopular!
- But just as in the previous examples, a gas is formed as a product of this reaction.

$$Na_2S\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + H_2S\ (g)$$

Stoichiometry of Precipitation

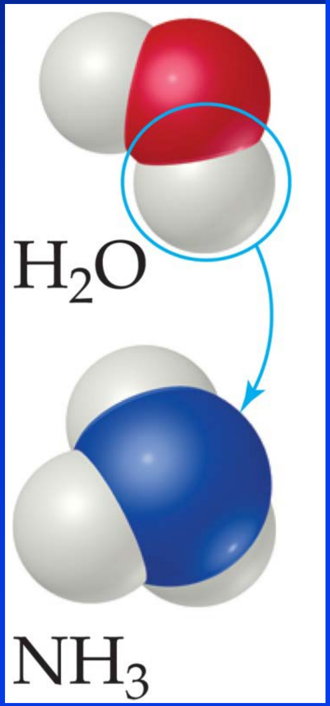
- Exactly the same, except you may have to figure out what the pieces are.
- What mass of solid is formed when 100.00 mL of 0.100 M Barium chloride is mixed with 100.00 mL of 0.100 M sodium hydroxide?
- What volume of 0.204 M HCl is needed to precipitate the silver from 50.ml of 0.0500 M silver nitrate solution ?

Types of Reactions

2 Acid-Base

- For our purposes an acid is a proton donor.
- a base is a proton acceptor usually OH^-
- What is the net ionic equation for the reaction of HCl(aq) and KOH(aq)?
- Acid + Base $\rightarrow$ salt + water
- $H^+ + OH^- \rightarrow H_2O$

Acid-Base Reactions



In an acid-base reaction, the acid donates a proton (H^+) to the base.

Neutralization Reactions

Generally, when solutions of an acid and a base are combined, the products are a salt and water.

$$CH_3COOH\ (aq) + NaOH\ (aq) \longrightarrow CH_3COONa\ (aq) + H_2O\ (l)$$

Neutralization Reactions

When a strong acid reacts with a strong base, the net ionic equation is…

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$$

Neutralization Reactions

When a strong acid reacts with a strong base, the net ionic equation is…

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$$

$$H^+\ (aq) + \cancel{Cl^-}\ (aq) + \cancel{Na^+}\ (aq) + OH^-(aq) \longrightarrow$$
$$\cancel{Na^+}\ (aq) + \cancel{Cl^-}\ (aq) + H_2O\ (l)$$

Neutralization Reactions

When a strong acid reacts with a strong base, the net ionic equation is…

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$$

$$H^+\ (aq) + \cancel{Cl^-}\ (aq) + \cancel{Na^+}\ (aq) + OH^-(aq) \longrightarrow$$
$$\cancel{Na^+}\ (aq) + \cancel{Cl^-}\ (aq) + H_2O\ (l)$$

$$H^+\ (aq) + OH^-\ (aq) \longrightarrow H_2O\ (l)$$

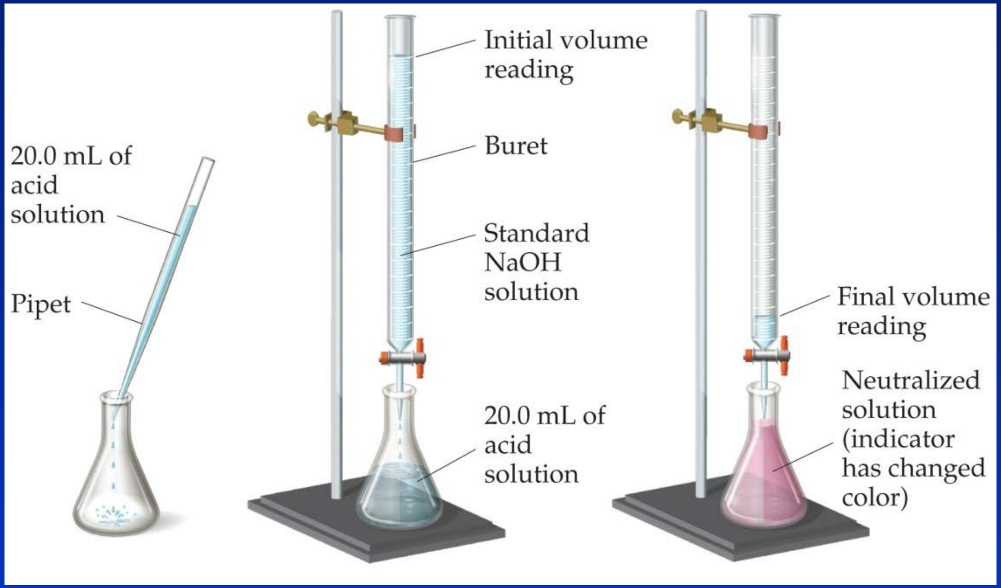
Acid - Base Reactions

- Often called a neutralization reaction Because the acid neutralizes the base.
- Often titrate to determine concentrations.
- Solution of known concentration (titrant), is added to the unknown (analyte), until the equivalence point is reached where enough titrant has been added to neutralize it.

Titration

- Where the indicator changes color is the endpoint.
- Not always at the equivalence point.
- A 50.00 mL sample of aqueous $Ca(OH)_2$ requires 34.66 mL of 0.0980 M Nitric acid for neutralization. What is $[Ca(OH)_2]$?
- # of H^+ x M_A x V_A = # of OH^- x M_B x V_B

Titration



Titration is an analytical technique in which one can calculate the concentration of a solute in a solution.

Acid-Base Reaction

- 75 mL of 0.25M HCl is mixed with 225 mL of 0.055 M $Ba(OH)_2$. What is the concentration of the excess H^+ or OH^- ?

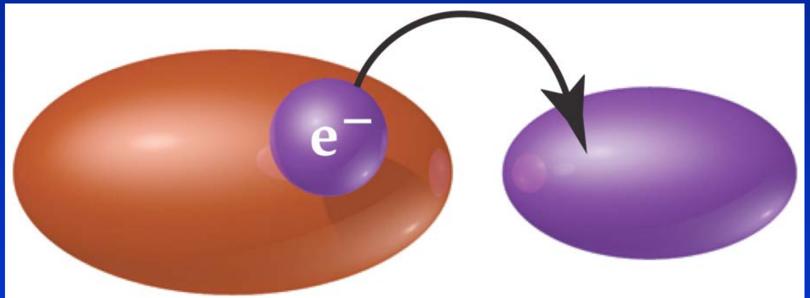
Types of Reaction

3 Oxidation-Reduction called Redox

- Ionic compounds are formed through the transfer of electrons.
- An Oxidation-reduction reaction involves the transfer of electrons.
- We need a way of keeping track.

Oxidation-Reduction Reactions

- An oxidation occurs when an atom or ion *loses* electrons.
- A reduction occurs when an atom or ion *gains* electrons.
- One cannot occur without the other.
- LEO says GER



Oxidation and Reduction

Chemical reactions often involve changes in oxidation states of atoms

Oxidation states are represented by oxidation numbers – these are essentially equal to the charge on a monatomic ion

$$Na \rightarrow Na^{+} + e^{-}$$

Sodium has changed its oxidation state from 0 to +1

EOS

Oxidation Numbers

- A way of keeping track of the electrons.
- Not necessarily true of what is in nature, but it works.
- MEMORIZE the rules for assigning
  - You must know these!

Determining Oxidation Numbers

Elemental form	**zero** (0). Only one kind of atom present, no charge
Atomic ions	**= the charge on the atom** (monatomic ion)
Group 1A **Li, Na, K, Rb, Cs**	**+1** unless in elemental form
Group 2A **Be, Mg, Ca, Sr, Ba**	**+2** unless in elemental form
Hydrogen (H)	**+1** when bonded to a nonmetal, **-1** when bonded to a metal
Oxygen (O)	**-1** in peroxides O_2^{-}, **-2** in all other compounds (most common)
Fluorine (F)	**-1**, always
Neutral compounds	The sum of all oxidation numbers of atoms or ions in a neutral compound is **zero**.
Ionic compounds	The sum of all oxidation numbers of atoms in an ionic compound is the **charge** on the polyatomic ion.

Oxidation States

Ex. 1

+1 because it's bound to a non-metal

-1 because it's an atomic ion with a -1 charge

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

0 because this is the pure elemental form of H

0 because it's an element in its pure form

+2

-2

These are just the ionic charges

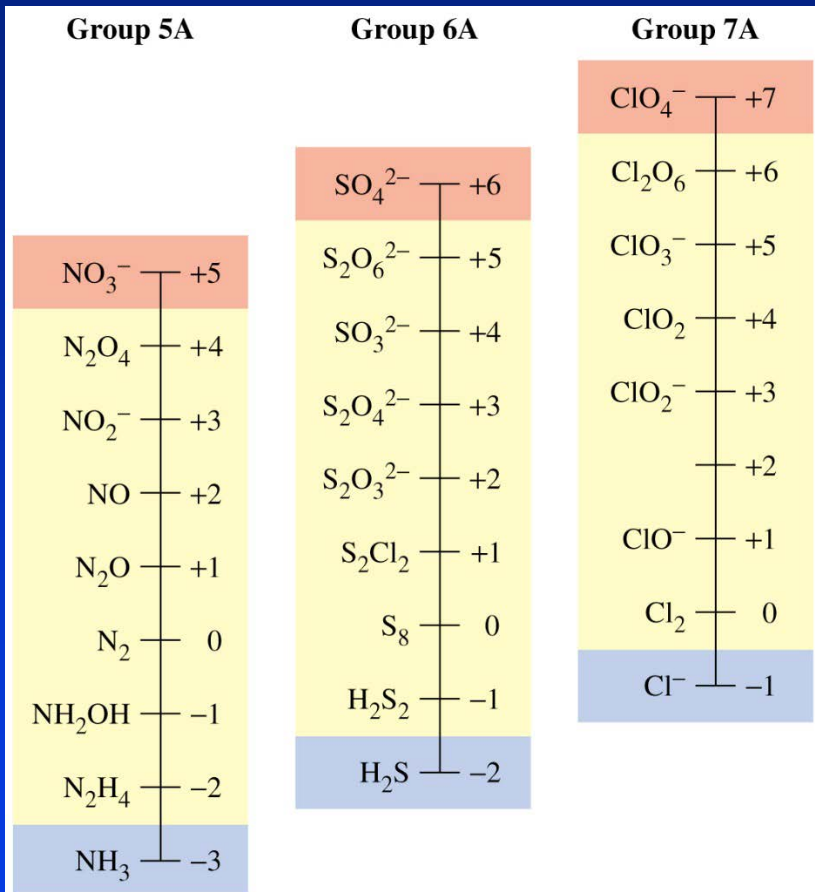
Ex. 2

Reduction

$$\overset{+4\ -2}{MnO_2} + \overset{+1\ -1}{4HCl} \longrightarrow \overset{+2\ -1}{MnCl_2} + \overset{0}{Cl_2} + \overset{+1\ -2}{2H_2O}$$

Oxidation

Oxidation Numbers of Nonmetals

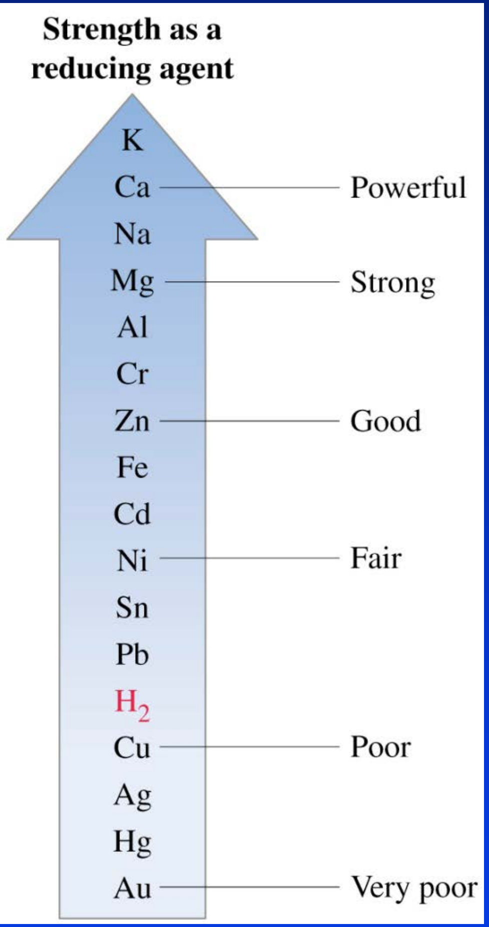


Max. ox. # = Group #

Min. ox. # = Group # – 8

Species in the middle of their oxidation number range can act as *either* oxidizing or reducing agents

Metals as Reducing Agents



Activity series: a metal will displace of any metal that lies **below** it in the series

e.g., Zn will replace H_2

$$Zn + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Oxidation States

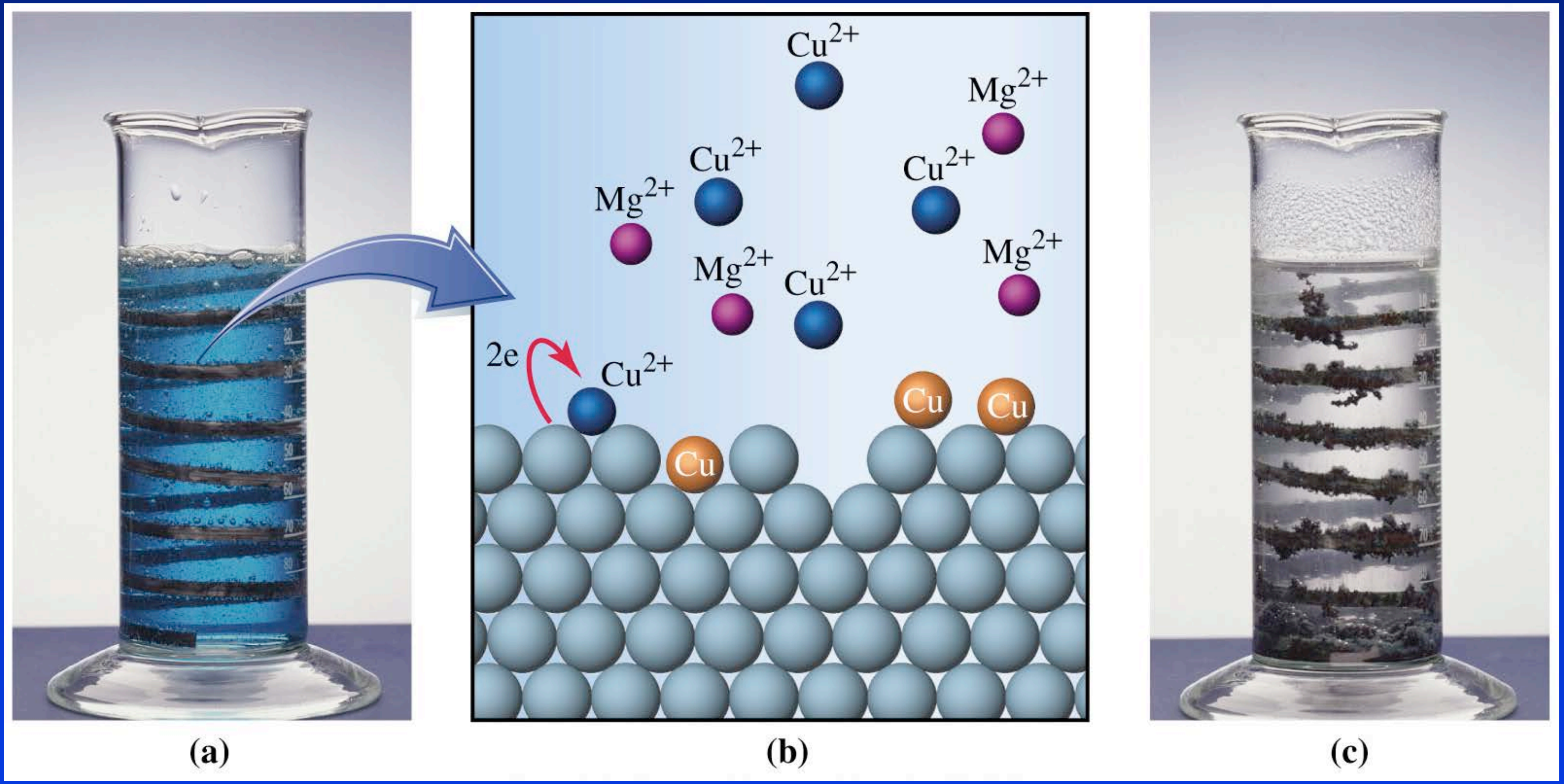
- Assign the oxidation states to each element in the following.
- CO_2
- NO_3^-
- H_2SO_4
- Fe_2O_3
- Fe_3O_4

Oxidation-Reduction

- Oxidation = increase in oxidation state
- Reduction = decrease in oxidation state

- The substance that is oxidized is called the reducing agent.
- The substance that is reduced is called the oxidizing agent.

An Example Redox Reaction



Agents

- Oxidizing agent gets reduced.
  - Gains electrons.
  - More negative oxidation state.
- Reducing agent gets oxidized.
  - Loses electrons.
  - More positive oxidation state.

Identify the in the following rxns

- Ox agent
- Red agent
- Substance oxidized
- Substance reduced

- $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$
- $Fe_2O_{3(s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(l)} + 3\ CO_{2(g)}$
- $SO_3^{2-} + H^+ + MnO_4^- \rightarrow SO_4^{2-} + H_2O + Mn^{2+}$

Writing and Balancing Redox Equations

- Must balance both mass and electric charge
- In water reactions, water can be either a reactant or a product
- In acidic reactions, H^+ can be either a reactant or a product
- In basic reactions, OH^- can be either a reactant or a product
- In **disproportionation reactions**, a portion of the reactant is oxidized and a portion of that same reactant is reduced

Half-Reactions

- All redox reactions can be thought of as happening in two halves.
- One produces electrons - Oxidation half.
- The other requires electrons - Reduction half.
- Write the half reactions for the following.
- $Na + Cl_2 \rightarrow Na^+ + Cl^-$
- $SO_3^{2-} + H^+ + MnO_4^- \rightarrow SO_4^{2-} + H_2O + Mn^{+2}$

Balancing Redox Equations

- In aqueous solutions the key is the number of electrons produced must be the same as those required.
- For reactions in acidic solution an 8 step procedure.

1 Write separate half reactions

2 For each half reaction balance all reactants except H and O

3 Balance O using H_2O

Acidic Solution

4 Balance H using H^+

5 Balance charge using e^-

6 Multiply equations to make electrons equal

7 Add equations and cancel identical species

8 Check that charges and elements are balanced.

Practice

- The following reactions occur in aqueous solution. Balance them
- $MnO_4^- + Fe^{+2} \rightarrow Mn^{+2} + Fe^{+3}$
- $Cu + NO_3^- \rightarrow Cu^{+2} + NO(g)$
- $Pb + PbO_2 + SO_4^{-2} \rightarrow PbSO_4$
- $Mn^{+2} + NaBiO_3 \rightarrow Bi^{+3} + MnO_4^-$

Now for a tough one

- $Fe(CN)_6^{-4} + MnO_4^- \rightarrow$
  $Mn^{+2} + Fe^{+3} + CO_2 + NO_3^-$

Basic Solution

- Do everything you would with acid, but add one more step.
- Add enough OH^- to both sides to neutralize the H^+
- Makes water
- $CrI_3 + Cl_2 \rightarrow CrO_4^{2-} + IO_4^- + Cl^-$
- $Fe(OH)_2 + H_2O_2 \rightarrow Fe(OH)^-$
- $Cr(OH)_3 + OCl^- + OH^- \rightarrow$
  $CrO_4^{2-} + Cl^- + H_2O$

Redox Titrations

- Same as any other titration.
- The permanganate ion is used often because it is its own indicator. MnO_4^- is purple, Mn^{+2} is colorless. When reaction solution remains clear, MnO_4^- is gone.
- Chromate ion is also useful, but color change, orangish yellow to green, is harder to detect.

Example

- The iron content of iron ore can be determined by titration with standard $KMnO_4$ solution. The iron ore is dissolved in excess HCl, and the iron reduced to Fe^{+2} ions. This solution is then titrated with $KMnO_4$ solution, producing Fe^{+3} and Mn^{+2} ions in acidic solution. If it requires 41.95 mL of 0.205 M $KMnO_4$ to titrate a solution made with 0.6128 g of iron ore, what percent of the ore was iron?